pH and pKa calculations

$$pH = - \log [H_3O^+]$$

Calculate the pH of 10mL of 0.015M HNO_3 if 20mL of water is added to it.

The concentration of diluted HNO_3 is $1/3^{rd}$ of $0.015 = 0.015/3 = 5 \times 10^{-3}$

$pH = - \log [5 \times 10^{-3}] = 2.30$

$$K_w = 1 \times 10^{-14} = [H_3O^+] [OH^-]$$

Calculate [OH-] of the diluted HNO_3 solution

$[H_3O^+]$ is the concentration in $molL^{-1}$ of the strong acid

$$[OH^-] = \frac{1 \times 10^{-14}}{5 \times 10^{-3}} = 2.0 \times 10^{-12}$$

$$K_a = \frac{[H_3O^+] [A^-]}{[HA]}$$

$$pK_a = - \log K_a$$

Calculate the pH of 20mL of 0.05M ethanoic acid if 20mL of water is added to it.

$K_a(CH_3COOH) = 1.78 \times 10^{-5}$

Ka value describes the extent that the acid will react with water. CH_3COOH is a weak acid so the K_a value is low, indicating that it is only slightly reactive with water so the Ka value must be used in the calculation

The concentration of diluted ethanoic acid is 1/2 of $0.05 = 0.05/3 = 0.025 molL^{-1}$

$$\mathbf{CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+}$$

$$Ka = \frac{[CH_3COO^-] [H_3O^+]}{[CH_3COOH]}$$

$$1.78 \times 10^{-5} = \frac{x \cdot x}{0.025}$$

$\sqrt{}$ (1.78 x 10-5) (0.025) = x

$6.67 \times 10^{-4} = x$ (this is the concentration of H_3O^+)

$pH = - \log (6.67 \times 10^{-4})$

pH = 3.18

$$K_b = \frac{[H_3O^+] [A^-]}{[HA]}$$

$$K_b = \frac{K_w}{K_a}$$

Calculate the pH of a 0.55M solution of ammonia

$K_a (NH_4^+) = 5.75 \times 10^{-10}$

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{5.75 \times 10^{-10}}$$

$$\mathbf{NH_3 \rightleftharpoons NH_4^+ + OH^-}$$

$$K_b = \frac{[NH_4^+] [OH^-]}{[NH_3]}$$

$$1.74 \times 10^{-5} = \frac{x \cdot x}{0.55}$$

$\sqrt{}$ $(1.74 \times 10^{-5}) (0.55) = x$

$3.09 \times 10^{-3} = x$ (this is the concentration of OH^-)

$$[H_3O^+] = \frac{1 \times 10^{-14}}{3.09 \times 10^{-3}}$$

$[H_3O^+] = 3.24 \times 10^{-12}$

pH = - log 3.24×10^{-12} = **11.5**